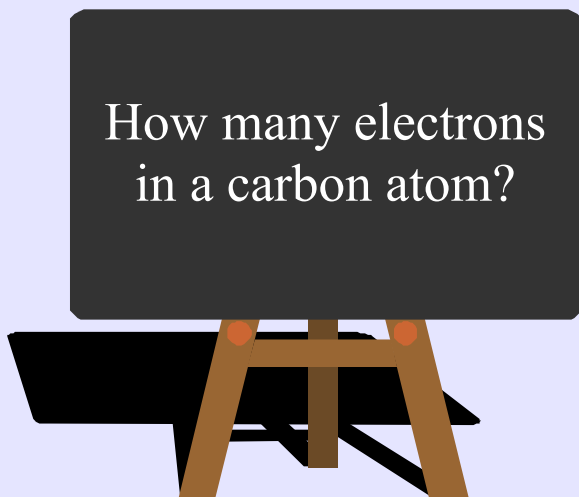


Chemistry Tests

Questions

How many electrons
in a carbon atom?



Dr David's Chemistry Tests (Questions)

Test 1: Theme: Atomic Structure.

1. Atomic number is the number of in the nucleus of an atom.
2. Electrons carry a electrical charge and are approximately the mass of a proton.
3. The isotopes of an element differ only in the number of in their nuclei.
4. A sodium atom has 12 neutrons in its nucleus and a mass number of 23. What are the values of x and y in the symbol, ${}^x_y\text{Na}$?
5. A line spectrum provides evidence for the movement of between the of an atom.
6. An orbital is a region of space about the nucleus of an atom in which there is greatest probability of finding a particular electron. Illustrate the shape of, (i) an s orbital and (ii) a p orbital.
7. In the mass spectrometer a powerful magnetic field is used to deflect charged particles through a semi-circular trajectory onto the plate. These particles are separated and sorted according to their
8. Define first ionisation enthalpy (energy).
.....
.....
9. Write a simple equation to represent the first electron affinity of chlorine.
.....
- 10 Calcium carbonate has the formula, CaCO_3
(i) Calculate the relative molecular mass (ie, formula mass) of this compound.
(RAM's: Ca - 40, C = 12, O = 16)
(ii) What mass of calcium carbonate will undergo complete reaction with 25 cm^3 of 0.1M HCl?

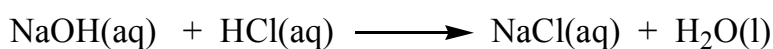


Test 2: Themes: Atomic Structure and the Mole.

1. An element may be represented, $^{31}_{15}\text{P}$. State, (a) the atomic number, (b) the mass number and (c) the number of neutrons in an atom of the element.

2. State the number of, (i) electrons, (ii) protons and (iii) neutrons in a calcium ion (Ca^{2+}).

3. 25.00 cm^3 of 0.1M sodium hydroxide solution neutralises 23.50 cm^3 of hydrochloric acid. Calculate the molarity of the acid.



4. Explain the importance of the *risk assessment form* in practical chemistry.

5. Why is it important to rinse out a pipette with the reagent you are going to put in it before using it?

6. What flame colour do calcium salts give when heated in a bunsen flame?

7. How many isotopes are there of chlorine? In what proportion do they occur in the natural element?

8. What is an orbital? Sketch the shapes of s and p orbitals.

9. In mass spectrometry, what is the purpose of the magnetic field?

10. Define the *mole*.

How much is a mole of sodium in terms of, (i) atoms of sodium and (ii) mass of sodium?

Test 3: Theme: Atomic Structure, Formulae.

1. Complete the electronic structure by replacing the asterisks with numbers. Name the element the structure corresponds to.

$1s^2, 2s^2, 2p^*, 3s^*, 3p^5$.

2. State the number of, (i) electrons, (ii) protons and (iii) neutrons in a fluoride ion (F^- , mass number = 19, atomic number = 9).

3. Write an equation for the decomposition of hydrated copper sulphate when it is heated by a bunsen flame.

State what you would *observe* when the product of the decomposition reacts with water.

4. Describe a test for sodium (either the element or a compound of the element).

5. Define the term, *relative isotopic mass*.

6. In the sodium atom how many electrons are in p orbitals?

7. Write the chemical formulae of the following:

Sodium carbonate,

Aluminium nitrate,

Potassium chloride,

Sulphuric acid.

8. Define, first ionisation energy by writing an equation.

Why does first ionisation energy increase across a period?

9. In mass spectrometry, what is the purpose of the *ionisation chamber*?

10. Define, *relative atomic mass*.

How many grams is (i) 0.1 mole of sodium, (ii) 1.5 mole of water?

Test 4: Theme: Atomic Structure, Formulae, Moles.

1. Copper has the atomic number 29. Write its electronic structure in s, p, d notation.
2. Name one element with diatomic molecules and one with triatomic molecules.
3. Describe a test for carbon dioxide gas. Name a compound which produces carbon dioxide when heated.
4. Describe a test for potassium (either the element or a compound of the element).
5. Define the term, *molarity*.
6. There are two main types of chemical bond, name them.
7. Write the chemical formulae for the following:
Nitric acid:
Sodium sulphate:
Trichloromethane (chloroform):
Calcium chloride:
8. Write an equation corresponding to *first electron affinity*.
9. Write the *empirical* formula of butane (molecular formula, C_4H_{10}).

Calculate the % carbon and hydrogen in methane (CH_4).
10. Write the ionic equation for the reaction of zinc with hydrochloric acid.

Test 5: Theme: Formulae & Bonding.

1. The aluminium ion is highly polarising. Explain what this means and why it is highly polarising.
2. Draw a dot-&-cross diagram to illustrate the bonding in sodium chloride.
3. In a crystal of sodium chloride, illustrate how the particles are arranged.
4. The water molecule is a polar molecule. What does this mean? Illustrate hydrogen bonding in water.
5. Sketch and explain the shapes of the following molecules: (a) ammonia, (b) sulphur dioxide.
6. How does a pi bond differ from a sigma bond?
7. Define relative atomic mass.
8. Write the ionic equation for the reaction of sodium hydroxide with hydrochloric acid.
9. 0.200 moles of an oxide of potassium has a mass of 22.0 g. Its empirical formula is KO, deduce its molecular formula.
10. Write the formula of phosphoric acid.

Test 6: Theme: Periods, Groups , Bonding & Equations.

1. Sodium and magnesium are s-block elements and they are in period 3 of the periodic table. Why are they referred to as s-block elements?

Which element in period 3 has an amphoteric character?

Name one element in period 3 with a basic oxide and one which has an acidic oxide.

2. Name an element in group 1 which has a superoxide.

Write the formula of this oxide and show how it reacts with water.

3. Why is the boiling point of sulphur higher than that of chlorine?

4. Sketch the structures of graphite and diamond.

5. By drawing dot-&-cross diagrams show the mechanism of bonding in calcium oxide.

6. How do the first ionisation energies of the elements Be to Ba vary? Explain the trend.

7. Write an ionic half equation for iron(II) being oxidised to iron(III).

Write an ionic half equation for chlorine being reduced to chloride ions.

Write the overall equation for the reaction between chlorine and iron(II) ions.

Test 7: Theme: Structure & Bonding.

1. Using dot-&-cross diagrams describe and illustrate the bonding in potassium chloride (atomic numbers: K = 19, Cl = 17).

2. List the general properties of covalent compounds.

3. Define, or explain, the term, *electronegativity*.

4. The polarising power of the aluminium ion is high compared with that of the sodium ion. What does the term, polarising power mean (or refer to)?

What is the effect of polarising power on the physical properties of aluminium chloride?

5. Define the term, *first ionisation energy* (enthalpy).

6. Describe, briefly, the metallic structure of magnesium.

7. What type of intermolecular forces are present in a crystal of iodine. Explain, briefly, how they act.

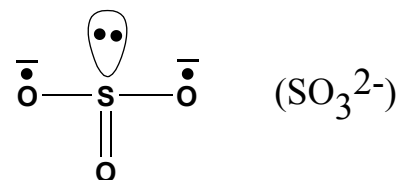
8. The water molecule is an angled molecule. Illustrate the shape of the molecule and explain why it has that shape.

9. Write the chemical formulae of: sulphuric acid, sodium carbonate, silver nitrate and barium chloride.

10. Give a chemical test for the sulphate ion.

Test 8: Theme: Redox, Structure, Bonding & Periodicity.

1. The chemical bonding in the sulphite ion may be represented:



Draw the shape of this ion and name the shape you depict.

2. Define oxidation in terms of electron transfer.

3. Iron reacts with chlorine, ie, $2\text{Fe(s)} + 3\text{Cl}_2\text{(g)} \rightarrow 2\text{FeCl}_3\text{(s)}$

Name the type of chemical bonding in FeCl_3 . In terms of electron transfer which element is being oxidised and which is being reduced? Explain.

Oxidised:

Reduced:

4. State the oxidation number of an atom of the underlined element in the following:-

(i) $\underline{\text{Mn}}\text{O}_4^-$

(ii) $\underline{\text{C}}\text{H}_4$

(iii) $\underline{\text{Al}}_2\text{O}_3$

5. On warming, the chlorate ion, $\text{OCl}^-(\text{aq})$, undergoes reaction to give Cl^- and ClO_3^- .

ie, $3\text{OCl}^-(\text{aq}) \rightarrow 2\text{Cl}^-(\text{aq}) + \text{ClO}_3^-(\text{aq})$

.....

Write the oxidation numbers of each atom on the dotted lines.

Considering any changes in oxidation numbers state the *type* of reaction. Explain.

6. Considering the s-block elements state the general trends in the following properties.

(i) change in atomic radii down the groups.

(ii) change in first ionisation energy down the groups.

(iii) change in thermal stability of the nitrates from group 1 to group 2.

(iv) change in polarising power of the cations down group 1.

7. What is it about the metallic structure of magnesium that gives it a higher melting point than sodium?

8. State the flame colours of, potassium, copper and calcium.

9. Use a dot & cross diagram to depict the chemical bonding in carbon dioxide.

10. What are the common oxidation numbers of the following elements in their compounds?

(i) hydrogen, (ii) oxygen, (iii) chlorine, (iv) sodium.

Test 9: Theme: Periodic Table.

Quick Fire Questions

1. Why do fires burn?
2. State a test for oxygen.
3. What colour is chlorine gas?
4. In the First World War what was chlorine gas used for?
5. State one difference between hydrogen and helium.
6. Is sulphur a metal or a non-metal?
7. Which has the highest electrical conductivity, aluminium or sulphur?
8. Is mercury a metal or non-metal?
9. Which metal has the highest melting temperature?
10. What nationality was Mendeleev?
11. What was Mendeleev's main scientific interest?
12. The vertical columns in the periodic table are called what?
13. What do elements in the same vertical column have in common?
14. What important property increases as the vertical arrangement of the noble (inert) gases is descended?
15. What is the family name of the group I elements?
16. Are the group one elements soft or hard?
17. How can you prevent the corrosion of sodium?
18. Which of the group I elements reacts most vigorously with water?
19. Write a balanced equation for the reaction of sodium with water.
20. What colour is the potassium flame?
21. How many atoms of fluorine combine with one atom of silicon?
22. How many atoms of fluorine combine with one atom of chlorine?
23. What happens to aluminium powder when tipped into a bunsen flame?
24. Write a formula for a phosphorus oxide.
26. State the appearance of silicon.
27. State two forms of carbon.
28. Are diamonds for ever?
29. Can diamonds be made to react with oxygen?
30. State one physical difference between silicon dioxide and carbon dioxide.
31. Why did Mendeleev put silicon and carbon in the same vertical group?
32. Which reacts most readily with water, lithium or sodium?
33. Do copper, silver and gold react with water?
34. On which side of the periodic table do metals occur?
35. Where are semi metals found in the periodic table?

Test 10. Theme: Halogens

1. State how the reactivity of the halogens changes as the group is descended.
2. State how the atom sizes of the halogens change as the group is descended.
3. State how the attraction of the nucleus for the outer electrons, in the halogen atoms, changes as the group is descended. Explain the change.
4. Arrange the halogens in order of electronegativity (the most electronegative first).
5. State what would be observed if chlorine gas was bubbled into a solution of potassium bromide. Explain the result.
6. State what would be observed if bromine was added to a solution of potassium iodide. Explain.
7. Which of the following halogens is the strongest oxidising agent: chlorine, bromine, iodine. Explain.
8. Describe and explain what would be observed if a solution of sodium bromide was added to a mixture of silver nitrate and nitric acid.
9. Give an example of a disproportionation reaction and explain what the term means.
10. Name a radioactive halogen.

Test 11. Theme: Sulphuric Acid.

1. Name a yellow coloured element used for the manufacture of sulphuric acid.
2. List the five main components of a sulphuric acid plant.

(i)	(ii)	(iii)	(iv)	(v)
-----	------	-------	------	-----
3. When sulphur burns in air what is the colour of the flame?
4. Name the gas formed when sulphur burns in air.
5. Name the catalyst used in the main reaction chamber of the contact process.
6. Write the equation for the reaction used in the main reaction chamber.
7. Is the above reaction exothermic or endothermic?
8. What is the best temperature for oxidising sulphur dioxide to sulphur trioxide?
9. To what is sulphur trioxide added to produce 98.5% sulphuric acid?
10. Why is sulphur trioxide not combined directly with water?

Write an equation for the reaction of sulphur trioxide with water.

11. Roughly how much sulphuric acid is produced, per hour, from your average sulphuric acid plant?
12. Name three products made from sulphuric acid.
13. What is oleum?
14. Round about 1900 the new contact process, for making sulphuric acid, was competing with an older method of manufacture dating from about 1740. What was the name of this older process and what were the essential reactants?
15. When, do you think, sulphuric acid was first referred to in the scientific literature?

Test 12. Theme: General Chemistry.

1. Arrange the following in order of reactivity (most reactive first):

Na, Cs, Ca, Li, Mg, Be

2. State the flame colours of potassium, calcium and barium:

Potassium:..... Calcium:..... Barium:

3. What are the oxidation numbers of the alkali metals and the alkaline earth metals?

Alkali metals:

Alkaline earth metals:

4. The s-block elements react with oxygen. Write the formulae of the oxides which are formed when the following metals are heated in oxygen: sodium, lithium, potassium and calcium.

Na:..... Li: K:..... Ca:

Write a balanced equation to show how one of these oxides reacts with water.

.....

5. State how the solubilities of the sulphates and hydroxides vary down group 2.

Sulphates:.....

Hydroxides:.....

6. What are the physical states of the following elements at room temperature and pressure?

Chlorine:

Bromine:

Iodine:

7. Write a balanced equation to show how chlorine reacts when passed over heated iron filings. Name the product.

.....

8. Chlorine reacts with water according to the following equation:



State the oxidation numbers of chlorine in this equation.

ie, in, Cl_2 :

in, HOCl:

in, HCl:

What do you understand by the term, *disproportionation*?

State a test for chlorine.....

Test 13: Theme: Organic Chemistry I.

1. Define or explain the term, homologous series.
2. Write general formulae for the alkene and chloroalkane homologous series.
3. Write the formula for, 2,3-dichloro-2-methylbutane.
4. Write an equation for the reaction of cyclohexene with bromine.
5. A dilute solution of potassium permanganate can be used to test for unsaturation. Assuming a positive result, what observations are expected?
6. Draw two structural isomers with the molecular formula, $\text{C}_2\text{H}_4\text{Cl}_2$.
7. Draw and name a pair of geometrical isomers.
8. Write formulae to show how the haloalkane, 2-bromopropane can be made from propene.
9. Show how 2-bromopropane reacts with aqueous potassium hydroxide. Name the product.
10. Write formulae to illustrate how polyethene may be prepared.

Test 14: Theme: Organic Chemistry.

1. Write a balanced chemical equation for the combustion of propane.
2. Write the formula of a chlorine free radical showing all of its outer electrons. How may these radicals be produced from chlorine molecules?
3. What is the term which is used to refer to the typical reactions of alkenes?
4. State one important commercial application of the hydrogenation of alkenes.
5. What features must an alkene molecule have in order to exhibit geometrical isomerism?
6. By writing formulae show how bromoethane can be converted into propionic acid ($\text{C}_2\text{H}_5\text{COOH}$).
7. Write the structural formula of butan-2-ol. What class of alcohol is this? Show how it reacts with phosphorus pentachloride (PCl_5). Name the organic product.
8. Both primary alcohols and secondary alcohols react with potassium dichromate solution acidified with sulphuric acid. However, the type of products are different. Explain this and give examples.
9. Write an equation showing how ethene reacts with bromine. Name the product and write the formula of the electrophile involved in this reaction.
10. What do you understand by the terms homolytic and heterolytic fission? Give examples to illustrate.

Test 15: Theme: Chemical Equilibrium.

1. When we say that a chemical reaction is *reversible* what do we mean?
2. Give an example of a reaction which is *not* reversible.
3. Give an example of a reaction which *is* reversible.
4. When silver nitrate solution is added to iron(II) nitrate solution a grey-black precipitate is formed and after a few minutes there is no further change. The mixture is now in *dynamic equilibrium*.
Explain what this term means.
5. Sometimes we say that the equilibrium of a chemical reaction lies to the right. What does this say about the *yield* of the reaction?
6. When writing the equation for a reversible chemical reaction would it be appropriate to use an arrow, an equal sign or double arrows? Explain.
Write the equation for the reaction referred to in question 4 using the appropriate symbol between the reactants and the products.
7. For a reversible reaction which has attained equilibrium, sketch a graph which shows how the rate of reaction, with respect to both reactants and products, varied from the start to some time after equilibrium had been attained.
8. State Le Chatelier's principle.

Test 16: Theme: Thermochemistry.

1. Give one example of an *exothermic* reaction and one example of an *endothermic* reaction. State what the terms mean.
2. Define *standard enthalpy of formation*.

State one application of standard enthalpy of formation values.

3. Calculate the enthalpy of combustion of methanol given the following data.
Equation for combustion:



Enthalpies of formation:

$$\Delta H_{\text{f}} \text{CH}_3\text{OH} = -239 \text{ kJmol}^{-1}$$

$$\Delta H_{\text{f}} \text{CO}_2 = -393 \text{ kJmol}^{-1}$$

$$\Delta H_{\text{f}} \text{H}_2\text{O} = -286 \text{ kJmol}^{-1}$$

4. State *Hess's law*.

5. Carbon monoxide burns in oxygen to form carbon dioxide:



calculate the carbon-oxygen bond energy in carbon monoxide given the following data:

$$E(\text{C}=\text{O}) = +743 \text{ kJmol}^{-1}, E(\text{O}=\text{O}) = +496 \text{ kJmol}^{-1}$$

6. Arrange the following in order of bond energy: H-H, F-F, $\text{N} \equiv \text{N}$ (lowest bond energy first).
What is likely to be the major product when fluorine is added to a mixture of hydrogen and nitrogen?
7. The specific heat capacity of water is 4.2 J. Define this quantity.

Test 17: Theme: Thermochemistry & Organic Chemistry.

1. Draw an energy profile (energy level diagram) to show the relative energies of reactants, transition state and product(s) in an exothermic reaction.

2. Camping stoves are often fuelled by butane. A cylinder of butane may contain 2.5 kg of the liquid hydrocarbon.

(i) Define the standard enthalpy of combustion of butane.

(ii) Write a balanced chemical equation for the combustion of butane.

(iii) Calculate the number of moles of butane in a cylinder of the gas.

(iv) Given that the standard enthalpy of combustion of butane is -2877 kJmol^{-1} . Calculate the enthalpy change which would occur if all the butane in the cylinder was burned.

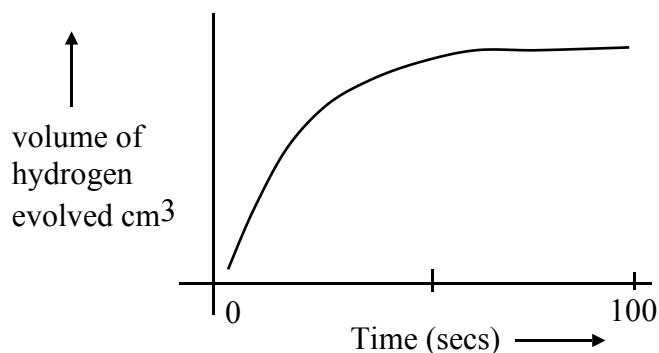
(v) Suggest two reasons why the enthalpy change you have just calculated is unlikely to be available to heat the vessel on the stove.

3. The butane molecule has a three dimensional structure. Show that you understand this statement by drawing a suitable structural formula.

Illustrate a structural isomer of butane. Name the structure illustrated.

Test 18: Theme: Chemical Kinetics.

1. Define **rate** of reaction.
2. State the factors which may be used to speed up a chemical reaction.
3. The rate of reaction of magnesium with a solution of hydrochloric acid can be investigated by introducing a piece of magnesium ribbon into excess 1M hydrochloric acid contained in a reaction flask connected to a graduated syringe. When plotted the results are as shown below.



- (i) Explain why the hydrogen is produced at a faster rate at the beginning of the experiment than at the end.
- (ii) Estimate how long it took for half of the magnesium to react.
- (iii) The experiment was repeated using a piece of magnesium ribbon twice the length of the first piece. The time taken for the ribbon to dissolve was roughly the same as on the first occasion. Explain.
- (iv) The experiment was repeated using the same conditions as in the first experiment except that a small amount of copper(II) sulphate solution was added to the reaction mixture to act as catalyst. What effect would this have on the results of the experiment?
- (v) If the experiment was repeated at a higher temperature (say 10° higher) the rate of reaction would increase, it may even double. Explain why raising the temperature a relatively small amount causes the rate to increase by a large amount. In your explanation refer to suitable Maxwell-Boltzmann distribution of molecular energies curves.

Test 19: Theme: Some Industrial Inorganic Chemistry.

1. Write an equation representing the Haber process for the manufacture of ammonia. State the manufacturing conditions. Explain how the conditions could be modified to provide a higher yield of ammonia. Explain why your suggestion is not adopted in practice.
2. How is nitric acid made from ammonia? State one important application of nitric acid.
3. Explain how chlorine and sodium chlorate(I) are manufactured. No diagrams are required although formulae and equations are required.
4. State important uses of chlorine and sodium chlorate(I).
5. Write a formula for sulphuric acid showing how the atoms are bonded to one another. Outline how sulphuric acid is manufactured stating the operating conditions. State six commercial products which require the use of sulphuric acid at some stage in their manufacture.

Test 20. Theme: Aluminium & Iron.

1. Aluminium is the most abundant metal in the earth's crust.

What percentage of the earth's crust is due to aluminium?

Give the name of the main ore of aluminium?

Write a formula for the principal aluminium compound in this ore.

Aluminium is manufactured by an electrolytic process. State:

- (i) the material used for the electrodes,
- (ii) the constituents of the electrolyte,
- (iii) the operating temperature of the electrolyte,
- (iv) the equation for the reaction at the anode and
- (v) the equation for the reaction at the cathode.

What is the principal requirement to make the manufacture of aluminium economic?

2. Iron is a transition element. The general properties of transition elements include; a number of oxidation states, coloured compounds, complex formation, catalytic activity. Give examples to show that iron exhibits these properties.

Give the name of an important ore of iron.

What is the formula of the principal chemical constituent of this ore?

Write a chemical equation to show how iron(III) oxide is reduced in the *blast furnace*.

Name the reducing agent.

How is this reducing agent produced in the blast furnace?

What is the chemical composition of *slag*?

What is the purpose of the *tuyeres* situated near the base of the blast furnace?

Name the type of iron run off from the base of the blast furnace?

State one important commercial application of iron.

Test 21: Theme: More Organic Chemistry.

1. The hydrolyses of haloalkanes provide examples of nucleophilic substitution. Write an equation representing the hydrolysis of 1-bromobutane in an aqueous solution of potassium hydroxide.

Write a formula for the nucleophile in this reaction.

State two ways in which the rate of this reaction could be increased.

Silver nitrate solution is sometimes used to test for haloalkanes. Explain how this works.

2. Show how pentanoic acid can be synthesised from 1-bromobutane.

3. Write a structural formula for cyclohexene.

What class of compound is cyclohexene?

How may cyclohexene be synthesised, in a one step process, from bromocyclohexane? Write formulae and state reaction conditions.

Are there likely to be any side products from this synthesis? Explain.

Explain the term, *elimination reaction*.

4. Name the following: $\text{CH}_2=\text{CHCH}_2\text{CH}_2\text{CHClCH}_2\text{CH}_3$

Test 22: Theme: More Organic Chemistry.

1. If 2-bromobutane was refluxed with an alcoholic solution of potassium hydroxide what would be the outcome?
2. How may 1-bromopropane be converted into propan-1-amine? Write an equation to illustrate and explain how the yield may be maximised.
3. What is DDT?

Since the 1940's the use of DDT has declined markedly. Explain why.

4. Alcohols are classified as primary, secondary tertiary, provide examples of each using alcohols containing four carbon atoms. Write the structural formulae and names of your examples.

5. Methanol has a relative molecular mass of 32 and its molecules contain only one carbon atom. It boils at 65°C . However, ethane which has a similar relative molecular mass (30) and contains two carbon atoms per molecule boils at -89°C . Explain this difference in boiling points.

6. Write a chemical equation for the combustion of propan-1-ol. Does it burn with a clean flame or a smokey dirty flame?

Test 23: Theme: More Organic Chemistry.

1. Explain how it is possible to distinguish between butan-1-ol and butan-2-ol using potassium dichromate acidified with dilute sulphuric acid.

2. How is it possible to make cyclohexene from cyclohexanol? Outline the experimental procedure.

What *type* of reaction is employed in the above preparation?

If the overall yield of cyclohexene is 50% how much cyclohexanol, in grams, should be used to make 2 grams of cyclohexene?

3. It is required to synthesise 1-chloropropane. Phosphorus pentachloride and a selection of alcohols are available. Outline a possible procedure. Write an equation for your proposed reaction.

4. A mixture of sodium bromide, concentrated sulphuric acid and butan-1-ol is refluxed. The apparatus is converted for distillation. What is the major organic product? Write chemical equations showing the course of the reactions.

Test 24: Theme: Some Environmental Chemistry.

1. Ozone is present in the earth's atmosphere. In which layer of the earth's atmosphere is there the highest concentration of ozone?

Why is this ozone layer beneficial?

Explain why the photodissociation reactions of ozone and oxygen in the atmosphere are important.

2. What are CFC's?

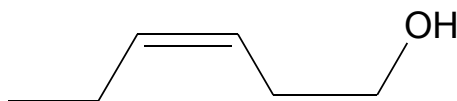
In 1930 the American chemist, Thomas Midgley, prepared dichlorodifluoromethane (Freon) and advocated its use as a refrigerant. Why did he consider this substance better than existing refrigerants?

It transpired that dichlorodifluoromethane and other related compounds deplete ozone in the atmosphere. How does this come about and what action can be taken to reduce the harmful effects of substances like dichlorodifluoromethane?

3. What is the *greenhouse effect*?

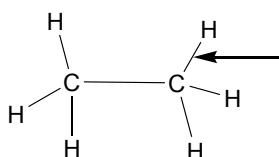
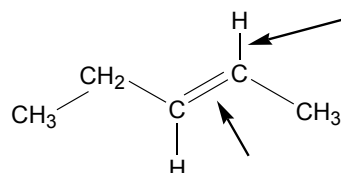
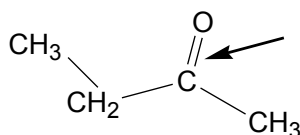
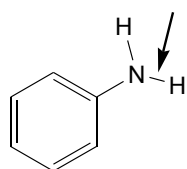
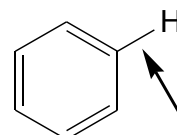
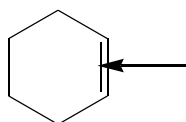
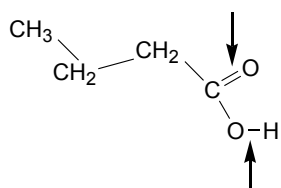
Why should we be concerned about the greenhouse effect?

4. Consider the following substance. It is known as *leaf alcohol* since it occurs in grasses, green leaves and herbs. What is its systematic name? If bromine was added across the double bond what would be the name of the product?



Test 25: Theme: Infra-red Spectroscopy.

1. When exposed to infra-red radiation chemical bonds vibrate. Taking the water molecule as an example, illustrate how its bonds vibrate under the influence of infra-red radiation.
2. Define the term, *wavenumber*.
3. On an infra-red spectrum where is the fingerprint region? Why is it called the fingerprint region?
4. If you suspected an unknown organic substance to be a saturated aldehyde, where would you look (ie, approximate wavenumber values) in its infra-red spectrum for confirmatory absorptions?
5. Describe the main, characteristic, features of the infra-red spectrum of ethanol.
6. Examine the formulae below and state the approximate ir wavenumber values for the bonds indicated.



Name each molecule.

Test 26: Theme: Stereoisomerism.

1. By illustrating but-2-ene explain what you understand by the term *geometrical isomerism*.
2. What are the essential structural features required for an alkene, such as but-2-ene, to exhibit geometrical isomerism?
3. By referring to the geometrical isomers of 3-chloropent-2-ene apply the Cahn-Ingold-Prelog system of nomenclature (ie, the *sequential rule*).
4. What is *plane polarised light*?
5. Butan-2-ol exists in two stereoisomeric forms known as enantiomers. Illustrate and explain this statement.
6. Draw structures to represent the optical isomers of 2-hydroxypropanoic acid (lactic acid).
7. Name the instrument used to measure the direction and angle of rotation of an optically active substance.
8. Explain the term, *racemic mixture*.

Test 27: Theme: More Organic Chemistry.

1. Write general formulae for aldehydes and for ketones.

2. How is Tollen's reagent made up? What is it used for? Explain.

3. When ethanal is made in the laboratory by oxidising ethanol, the ethanol and oxidising agent are added gradually to 30% sulphuric acid at about 50°C.

What is the oxidising agent of choice?

Why is the addition of reactants made in this way? Why not just mix them all together and distil?

Write a simplified equation for the oxidation.

4. How are propanone and benzenol (phenol) manufactured?

5. Sodium borohydride is a useful laboratory reducing agent. If it was used to reduce propanone what would you expect to be the major organic product?

What is the attacking nucleophile in this reaction?

6. The iodoform reaction is sometimes used as a test for ketones.

What is iodoform?

How is it used as a test reagent?

7. 2,4-dinitrophenylhydrazine is a test reagent for aldehydes and ketones.

Describe its application.

8. Write a general formula for a Grignard reagent.

By writing a suitable equation show how a named Grignard reagent can be used to synthesise propan-2-ol.

9. What are the main features of the IR spectrum of hexanal?

Would there be any major differences between the IR spectra of hexanal and hexan-2-one? Explain.

Test 28: Theme: Carboxylic Acids.

1. Draw a structural formula for butanoic acid.

Write a formula for 2-bromo-2-chloromethylpropanoic acid.

Ethanoic acid (also known as *acetic acid*) is freely soluble in water. It is present, for example, in vinegar to the extent of about 5%. Provide an explanation for its high solubility in water.

2. Ethanoic acid is referred to as a *weak acid*. Explain what you understand by the term, *weak acid*.

3. You are asked to characterise an unknown organic compound and having applied a few tests you suspected that it is a carboxylic acid. Give details of **two** tests you believe would confirm your suspicions.

4. How could propanoic acid be converted into butanoic acid? Write equations to illustrate a possible reaction sequence.

5. Benzenecarboxylic acid (*benzoic acid*) is a white crystalline solid melting at 121°C. Write a structural formula for the acid.

Describe briefly how you would measure the melting point of this acid.

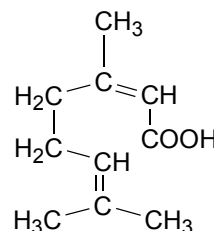
Benzenecarboxylic acid is sparingly soluble in cold water but freely soluble in cold dilute sodium hydroxide solution. Explain the difference.

How may benzenecarboxylic acid be converted into benzenecarboxamide (*benzamide*)? Write equations.

Test 29: Theme: Carboxylic Acids, Aromatics & Transesterification.

1. The infra-red spectrum of ethanoic acid shows a large, very broad, peak centred at about 3100 cm^{-1} and another large and fairly broad peak at 1720 cm^{-1} . Identify the bonding systems characteristic of these absorptions.

2. Give the common or systematic name of the following:



Where in nature has this acid been detected?

3. Draw a structural formula for benzene.

Why is the benzene molecule much more stable than molecules of aliphatic trienes?

What are the typical reactions of benzene?

4. Phenylethanone (acetophenone) can be synthesised from benzene using the Friedel-Crafts reaction. Write an equation and state the conditions for this reaction.

5. Write a formula for the organic substance obtained when benzenamine (aniline) is treated with an ice cold solution of nitric(III) acid (ie, nitrous acid; made by mixing ice cold solutions of sodium nitrate(III) and dilute hydrochloric acid).

With the aid of an equation show how an ice cold solution of this substance reacts with benzenol (phenol). Name the precipitated product.

6. Transesterifications are reactions in which one ester is converted into another. Explain how the manufacture of biodiesel provides a commercial example of transesterification.

Test 30: Theme: Terylene, Proteins and Amino Acids.

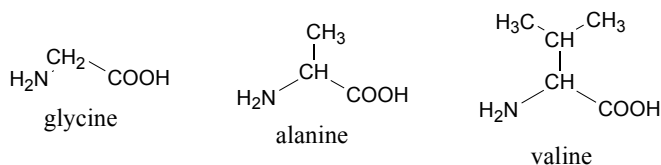
1. The polyester, Terylene, is synthesised by condensation polymerisation. Explain what you understand by the term, *condensation polymerisation*, and write an equation illustrating the formation of the polymer.

2. Proteins and carbohydrates are essential components of living material. How do proteins differ from carbohydrates in terms of chemical composition?

3. Write a general formula for an α -amino acid and explain why these molecules are important in relation to proteins.

4. Draw an α -amino acid with the S configuration.

5. Given the structures of glycine, alanine and valine show how they could join together to form a tripeptide.



6. The term, *dipolar ion* (or *zwitterion*) is used in relation to α -amino acids. Explain what this term means.

7. When a protein is boiled with hydrochloric acid a mixture of amino acids is formed. How would you set about identifying the amino acids in such a mixture?

Test 31: Theme: Acid-base Equilibria

1. Define pH.

2. Acid-base indicators are weak acids; their dissociation in aqueous solution may be represented:



HIn represents the indicator.

Bromophenol blue indicator has a K_{In} value of $1 \times 10^{-4} \text{ mol dm}^{-3}$. Calculate the pH of a solution in which $[\text{In}^-] = [\text{HIn}]$.

3. Calculate the pH of 0.02M Ba(OH)_2 assuming the base to be completely dissociated in aqueous solution.

4. Calculate the pH of 0.05M hydrochloric acid.

5. A buffer solution is made by mixing equal volumes of 1M ethanoic acid (acetic acid) and 1M sodium ethanoate (sodium acetate).

Write a chemical equation to show the chemical species which exist in equilibrium in this buffer solution.

Explain how this solution is able to maintain a pH of around 4.7 even when small quantities of hydrochloric acid or sodium hydroxide are added to it.

6. Sketch a graph showing how the pH of 25 cm³ of 0.1M ethanoic acid changes when titrated with 0.1M sodium hydroxide.

Label the buffer region.

Test 32: Theme: Chemical Equilibria.

1. Hydrogen reacts with iodine vapour to form gaseous hydrogen iodide in a reversible reaction. Write a balanced equation for the reaction.

Write an expression for the equilibrium constant, K_C .

Suppose this mixture is contained in a closed 2 dm^3 vessel at 600K . Calculate the value of the equilibrium constant, K_C , given that, $[\text{HI}] = 0.8 \text{ mol dm}^{-3}$, $[\text{H}_2] = 0.1 \text{ mol dm}^{-3}$ and $[\text{I}_2] = 0.1 \text{ mol dm}^{-3}$.

Does K_C have units? Explain.

2. Define, (i) mole fraction & (ii) partial pressure (p) in terms of mole fraction of reactant and total pressure (P_T).

3. Consider the following equilibrium: $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$

Write an expression for K_p for this equilibrium.

At a particular temperature K_p has a value of 2000 kPa . If the partial pressure of NO_2 is 200 kPa what is the partial pressure of N_2O_4 ? What is the total pressure exerted by the equilibrium mixture?

4. In each case give **one** example of homogeneous and heterogeneous equilibria.

5. At a certain temperature, ethanoic acid was allowed to equilibrate with ethanol. Analysis showed that the equilibrium mixture contained 2.0 mol dm^{-3} ethanoic acid and 3.5 mol dm^{-3} ethanol. The equilibrium constant, K_C , has a value of 4.0 at the temperature of the mixture. Write the equation for the equilibrium and calculate the concentrations of the other components of the mixture.

Test 33: Theme: Redox Equilibria.

1. An electrochemical cell containing a zinc half cell and a silver half cell was set up using a potassium nitrate salt bridge. The cell potential at 25°C was measured to be 1.40V.



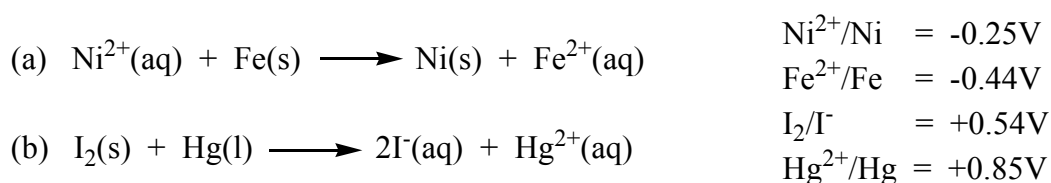
(a) Draw a labelled diagram of the cell.

(b) Calculate the standard cell potential.

(c) Suggest two reasons why the calculated value is different from the measured value.

(d) Write an equation for the overall cell reaction.

2. Using E^\ominus values, state which of the following reactions can take place. Give reasons.



3. Explain the difference between *primary* and *secondary* storage cells.

4. What do you understand by the term, *fuel cell*?

5. The standard electrode potential may be defined as the emf of a cell in which the left hand component is a standard hydrogen electrode (SHE) and the right hand component is the electrode being measured.

For zinc the standard electrode potential is -0.76V. What does the negative sign indicate? Which way would the electrons flow in such a cell? If the cell was to deliver current, what reactions would occur at the two electrodes and what would be the overall cell reaction?

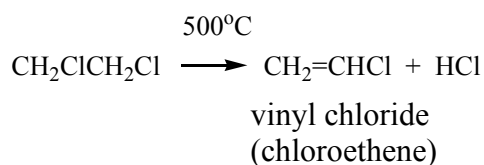
Test 34: Theme: Industrial Chemistry.

1. Write formulae for two sulphuric acids. Give their systematic names.
2. Explain what you understand by the term, *pickling*.
3. What is the chemical composition of the fertilizer referred to as *superphosphate*?
4. Chlorine is manufactured by the electrolysis of brine in the membrane or diaphragm cell. What is brine?

Write equations to represent the reactions at the cathode and anode in this cell.

Most chlorine is used to manufacture polyvinyl chloride (PVC). Write a general formula to represent this polymer.

A number of methods are available for making the monomer vinyl chloride. An important process involves the pyrolysis of 1,2-dichloroethane:



By writing an equation suggest a method for making 1,2-dichloroethane.

5. Write an equation for the *disproportionation* of sodium chlorate(I) in aqueous solution.

How is sodium chlorate(V) manufactured and what is it used for?

6. Name the main ore of aluminium. State six commercial applications of the metal.

7. Name and write the formula of a magnetic ore of iron.

If iron was to be extracted from this ore suggest a suitable reducing agent. Write an equation for the reaction.

8. Write a formula to represent the copper ore, *malachite*. What colour is this ore? Write an equation to for the thermal decomposition of the ore.

In the laboratory how would you produce a sample of copper from the thermal decomposition product?

9. Which has the highest melting point, titanium or tungsten?
10. Write a chemical equation for the reduction of tungsten(VI) oxide.

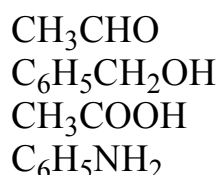
Test 35: Theme: Organic Analysis.

1. What do you understand by the term, *criterion of purity*?
2. The term *chromatography* literally means, colour writing. Give an example of an application of chromatography which illustrates its connection with colour.
3. Define the term, R_f value.
4. How does paper chromatography differ from thin layer chromatography (tlc)?
5. If you wish to separate two organic compounds using chromatography, on a gram scale, would you employ tlc or column chromatography? Explain your choice.
6. What do the letters HPLC stand for? Outline the main features of this technique.

Are there any advantages of HPLC over ordinary (gravity) column chromatography?

7. Outline the principles of gas liquid chromatography. Explain the terms, stationary phase, flame ionisation detector and retention time.

8. Consider the following molecules:
In each case name the substance and state the main characterising features in their ir and nmr spectra.



9. What is the purpose of the Lassaigne test?
10. Analysis of an unknown organic compound gives negatives for nitrogen, halogens and sulphur. It gives positives with Bradys and Tollens reagents. It also gives a positive iodoform test and it burns with a very smokey flame. What structural features do these tests indicate?

Test 36. Theme: Practical work - titrimetric analysis-1.

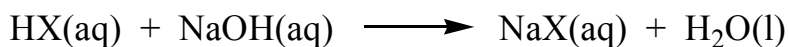
1. List all the apparatus required to perform a titration to standardise hydrochloric acid by titration against a standard solution of sodium hydroxide.
2. State the safety equipment and precautions required when performing an acid-base titration such as that referred to above.
3. Name the indicators you would recommend for the following titrations.
 - (i) 0.05M sulphuric acid versus 0.1M sodium hydroxide.
 - (ii) 0.1M sodium hydroxide versus 0.1M ethanoic acid.
 - (iii) 0.1M hydrochloric acid versus 0.1M ammonia solution.
4. Suggest a procedure, in the titration of sodium hydroxide against hydrochloric acid, which is likely to introduce a significant error unless it is carried out extremely carefully.
5. Make list of all the things you should do to minimise errors when using a burette to deliver a reagent during a titration.
6. Imagine you are titrating sodium hydroxide against hydrochloric acid. The indicator is changing colour and you have decided that you have reached the end point. On close inspection you observe that a drop of titrant is hanging from the tip of the burette. Which of the following would you advise?
 - (i) Ignore it.
 - (ii) Using a wash bottle of distilled water, wash the drop from the tip into the titration flask.
 - (iii) Leave it hanging from the tip and subtract 0.05 cm^3 from the burette reading.
 - (iv) Leave it hanging from the tip and add 0.05 cm^3 to the burette reading.
 - (v) Abort the titration and start again.
7. It is recommended that a white tile is placed under the titration flask. What is the purpose of this?
8. Sometimes during a titration the tip of the burette becomes blocked with grease or a particle of solid. What would you advise doing in such a circumstance?
9. Proper burette stands are usually made entirely of wood. Suggest a reason for this.
10. What should you do with the glassware after you have completed a titration exercise?

Test 37: Theme: Practical work - titrimetric analysis -2.

1. In an acid-base titration 23.25 cm^3 of sodium hydroxide was found to exactly neutralise 25.00 cm^3 of hydrochloric acid. The acid was 0.105M , what was the molarity of the alkali?
Write an equation for the reaction and work out the answer from first principles (ie, do not use a formula).
2. A standard solution of hydrochloric acid (0.105M) was titrated against 25.00 cm^3 of a dilute solution of sodium carbonate. Methyl orange was used as indicator and the end-point was recorded as 26.05 cm^3 .
Write an equation for the reaction and calculate the molarity of the sodium carbonate solution. Calculate also the concentration of the carbonate solution in grams per 250 cm^3
3. Exactly 1g of hydrated sodium carbonate was weighed out and dissolved in 250 cm^3 of distilled water in a graduated volumetric flask. 25.00 cm^3 aliquots of this carbonate solution were titrated against standard 0.10M hydrochloric acid using methyl orange indicator. Three titre values were obtained: 16.10 , 16.15 and 16.20 cm^3 .
Calculate the mean titre value. The formula of the hydrated sodium carbonate may be represented, $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$. Calculate the value of x in this formula.
In your opinion, why would phenolphthalein indicator have been unsuitable?
Explain how you would have transferred the sample of hydrated sodium carbonate to the graduated volumetric flask and made up the solution.

Test 38: Theme: Practical work - titrimetric analysis - 3.

1. 2.50 g of an unknown acid (HX) was weighed into a small beaker and dissolved in some distilled water. The solution was washed into a 250 cm³ volumetric flask and made up to the graduation mark with distilled water. The flask was shaken and 25 cm³ aliquots titrated against 0.10M NaOH using methyl orange indicator. The titre values were, 25.85, 25.95 and 25.90 cm³.



Calculate the relative molecular mass of the unknown acid (HX).

2.

2.50 g of pure, crystalline, calcium carbonate was weighed into a clean dry previously weighed beaker. About 25 cm³ of distilled water was added followed by exactly 25 cm³ of dilute nitric acid. The mixture was allowed to stand, with occasional swirling, at room temperature until effervescence had ceased.

After the reaction, the liquid was decanted off and discarded. The remaining crystals were rinsed a few times with distilled water and then allowed to dry in a fan assisted oven at about 40°C. The beaker was cooled and weighed again to ascertain the mass of calcium carbonate remaining.

Mass of calcium carbonate crystals initially = 2.50 g.

Mass of calcium carbonate crystals remaining = 1.15 g

Calculate the concentration of the dilute nitric acid in mol dm⁻³.

3. 10 cm³ of white wine vinegar was pipetted into 100 cm³ graduated volumetric flask. The solution was made up to volume with distilled water. The solution was shaken and 25 cm³ aliquots titrated against 0.10 M sodium hydroxide solution using phenolphthalein indicator. The mean titre value was 20.25 cm³.

Assuming that the only acid in the vinegar is ethanoic acid write an equation for its reaction with sodium hydroxide.

Calculate the molarity of the vinegar with respect to ethanoic acid. Also calculate the mass of ethanoic acid per 100 g of the white wine vinegar.

Test 39. Theme: Practical work - titrimetric analysis - 4.

1. State the colour of 0.02M potassium permanganate solution.

What mass of crystalline potassium permanganate (KMnO_4) must be dissolved in distilled water to make a litre of 0.02M solution?

Are there any precautions to take when using 0.02M potassium permanganate?

The titration of hydrochloric acid against sodium hydroxide is an example of an *acid-base* titration. What type of titrations involve potassium permanganate?

2. An approximately 0.02M solution of potassium permanganate was made up and standardised against iron(II) ammonium sulphate ($\text{FeSO}_4 (\text{NH}_4)_2 \text{SO}_4 \cdot 6\text{H}_2\text{O}$) solution. The iron(II) salt solution was made up using 9.80 g of high purity salt dissolved in 250 cm^3 of distilled water in a graduated volumetric flask.

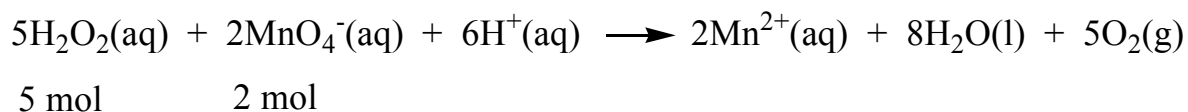
The burette was charged with the permanganate solution and 25 cm^3 of the iron(II) solution was pipetted into a 250 cm^3 titration flask. Approximately 25 cm^3 of 1M sulphuric acid was added to the titration flask and permanganate solution added, with swirling, until the solution in the flask acquired a permanent pink tinge. The titration was repeated a number of times and the mean titre value calculated as 23.40 cm^3 .

Write an ionic equation for the reaction and calculate the molarity of the permanganate solution.

Pure iron(II) ammonium sulphate can be used as a primary standard for standardising potassium permanganate. Explain the term, *primary standard* and state the general properties of a primary standard.

Test 40: Theme: Practical work - titrimetric analysis - 5

1. Hydrogen peroxide is a constituent of some mouthwashes. The hydrogen peroxide content may vary between 1% and 6%. Providing the mouthwash does not contain other reducing agents and is not highly coloured the hydrogen peroxide content can be established by titration against a standard solution of potassium permanganate.

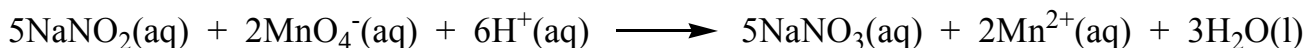


25.00 cm³ of a colourless hydrogen peroxide mouthwash was diluted to 250 cm³ with distilled water in a graduated 250 cm³ volumetric flask. The mixture was shaken and 25.00 cm³ aliquots, acidified with dilute sulphuric acid, titrated against standard 0.02M potassium permanganate solution. The mean permanganate titre was 31.25 cm³. Calculate the concentration of hydrogen peroxide in the original mouthwash. State the result in mol per dm³ and gram per dm³ and as a %.

What are the commercial applications of hydrogen peroxide?

2. Sodium nitrate(III) (sodium nitrite) is a whitish crystalline solid freely soluble in water. It is used commercially to make diazonium compounds and as a food preservative. With acids it decomposes giving nitrogen oxides. Its purity may be estimated by titration with standard potassium permanganate solution.

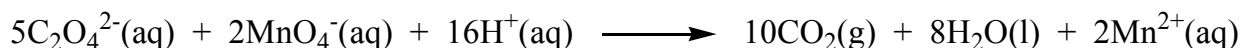
1.00 g of a sample of sodium nitrate(III) was dissolved in 250 cm³ of distilled water in a graduated volumetric flask. The solution was mixed thoroughly and some of it introduced into a burette. Potassium permanganate solution (25 cm³, 0.02M) was added to a shallow beaker and diluted with 15 cm³ of dilute sulphuric acid. The burette was positioned with its tip dipping into the permanganate solution and nitrate(III) solution added, with stirring, until the mixture in the beaker acquired a faint pink colour. The titration was repeated a number of times and the mean titre calculated. The mean titre of nitrate(III) solution was 24.85 cm³. Calculate the % purity of the sodium nitrate(III).



Why is the titration performed in the manner described with the tip of the burette dipping into the permanganate solution?

Test 41: Theme: Practical work - titrimetric analysis - 6.

1. A standard solution of potassium permanganate (0.02 M) was titrated against 25.00 cm³ of a solution of sodium oxalate. The mean titre value was 26.25 cm³. Calculate the molarity of the sodium oxalate solution.



If you were advising someone on how best to perform this titration in the laboratory what procedural details would you provide them with?

Are there any safety factors to consider?

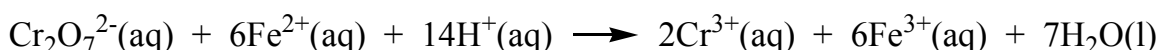
Oxalic acid can be used to standardise potassium permanganate. Does it have any advantages or disadvantages over its sodium salt?

2. What colour is potassium dichromate?

Calculate molar mass of potassium dichromate (K₂Cr₂O₇); you will need to look up the relative atomic masses of the elements.

1.226 g of pure potassium dichromate was dissolved in distilled water and the solution made up to exactly 250 cm³ in a graduated volumetric flask. Calculate the molarity of this solution.

This solution was introduced to a burette and titrated against 25.00 cm³ aliquots of an iron(II) solution. Each aliquot was acidified with 30 cm³ of 1M sulphuric acid and 5 cm³ of phosphoric acid. The end point was determined using 8 drops of barium diphenylamine sulphonate indicator (0.3% in water). The mean titre value was 26.85 cm³. Calculate the molarity of the iron(II) solution.



What is the purpose of the phosphoric acid?

What are the expected colour changes of the indicator during the titration?

Which is the strongest oxidising agent, permanganate or dichromate? Does dichromate have any advantages over permanganate in titrimetric analysis?

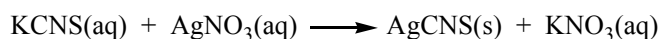
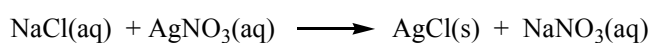
Test 42: Theme: Practical work - further titrations .

1. An impure sample of crystalline sodium chloride was weighed out and transferred to a 250 cm³ graduated volumetric flask. The flask was a quarter filled with distilled water, stoppered, and shaken to dissolve the crystals. The flask was opened, the stopper rinsed into the flask with a little distilled water, and exactly 150 cm³ (an excess) of 0.10 M silver nitrate solution added. Approximately 3 cm³ of concentrated nitric acid was added to the contents of the flask and the mixture swirled. The stopper was replaced and the mixture shaken to complete the precipitation of silver chloride. The flask was finally made up to the graduation mark with more distilled water and shaken to thoroughly mix the contents.

About half of the mixture, in the flask, was filtered into a clean beaker. Aliquots (50.00 cm³) of this filtrate were titrated against 0.10 M potassium thiocyanate (in the burette) using 3 cm³ of 10% iron(III) ammonium sulphate indicator solution.

The mass of the sample of impure sodium chloride was 0.450 g. The mean KCNS titre was 15.65 cm³. Calculate the % purity of the sodium chloride sample.

(This titrimetric method is attributed to the German chemist, Jacob Volhard (1834-1910), and is often referred to by his name)

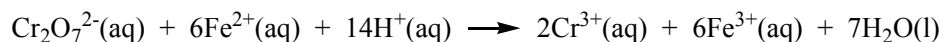


The old name for the indicator reagent is *ferric alum*. Discover the formula of this hydrated crystalline double salt. What colour does the indicator give in this titration?

Why is it that the indicator does not give a colour immediately the thiocyanate solution is added to the mixture in the titration flask?

2. Potentiometric titration.

A standard solution of potassium dichromate (0.017 M) was introduced into a burette and the level adjusted to read 0.00 cm³. Iron(II) sulphate solution (25.00 cm³) was added to approximately 25 cm³ of dilute sulphuric acid and 5 cm³ of phosphoric acid in a beaker. A redox probe, connected to a suitable voltmeter, was introduced into the solution in the beaker. Millivolt readings were recorded as small volumes of dichromate solution were added to the mixture in the beaker. The readings are shown below. Plot the results (mV versus cm³ of dichromate) and deduce the end point. Calculate the molarity of the iron(II) solution.



volume cm ³	millivolts
4.30	-300
9.15	-280
13.40	-260
18.20	-240
21.85	-210
24.30	-160
24.70	-60
24.80	-10
25.25	20
26.00	60
29.00	100
32.80	120
40.25	140
50.00	160

Test 43: Theme: Practical work - Inorganic Analysis - 1

1. Inorganic compound A was analysed with the following results. Fill in the inference column and name and write a name and formula for A.

TEST	OBSERVATION	INFERENCE
A small amount of A was taken up on a platinum wire and heated in a hot bunsen flame.	A bright yellow-orange colour was observed.	
A small amount of A was heated in a clean, dry, test tube	A colourless liquid condensed on the cool part of the tube.	
A small amount of A was dissolved in distilled water. Dilute hydrochloric acid was added.	No change was observed. No gas evolved.	
A small amount of A was dissolved in distilled water. The solution was acidified with dilute nitric acid and dilute silver nitrate added.	No change was observed.	
A small amount of A was dissolved in distilled water. The solution was acidified with dilute hydrochloric acid and dilute barium chloride added.	A white precipitate was obtained.	
	Name of A:	
	Formula of A:	

2.

TEST	OBSERVATION	INFERENCE
Performed flame test on pure substance B.	Gave a lilac coloured flame.	
The salt B was freely soluble in distilled water. A dilute solution was treated with dilute nitric acid and dilute silver nitrate.	A yellow precipitate was formed. This was insoluble in ammonia solution.	
An aqueous solution of B was treated with chlorine water.	The mixture acquired an immediate brown colouration. The solution turned starch paper blue.	
	Name of B:	
	Formula of B:	

Test 44: Theme: Practical work - Inorganic Analysis - 2

1.

TEST	OBSERVATION	INFERENCE
Pure substance C was heated strongly, on a platinum wire, in a bunsen flame.	An apple green colour was observed.	
Substance C is a white crystalline solid insoluble in water. However, it dissolves in dilute hydrochloric acid.	Efforvescence was observed. The gas evolved was colourless and odourless and turned lime water milky.	
A small amount of C was dissolved in dilute hydrochloric acid. 2 cm ³ of dilute sodium sulphate was added.	A white precipitate was formed which was insoluble in excess hydrochloric acid.	
	Name of C:	
	Formula of C:	

TEST	OBSERVATION	INFERENCE
Pure substance D was heated strongly, on a platinum wire, in a bunsen flame.	A brick-red colour was observed.	
Substance D is a white crystalline solid soluble in water. Dilute sulphuric acid was added to the solution.	A white precipitate was formed..	
A small amount of D was dissolved in distilled water and 2 cm ³ of dilute sodium carbonate was added.	A white precipitate was formed .	
A small amount of D was dissolved in distilled water. The solution was acidified with dilute nitric acid and 2 cm ³ of dilute silver nitrate added.	A white precipitate was formed which was soluble in dilute ammonia solution.	
	Name of D:	
	Formula of D:	

Test 45: Theme: Practical work - Inorganic Analysis - 3

1.

TEST	OBSERVATION	INFERENCE
Pure substance E was heated strongly, on a platinum wire, in a bunsen flame.	A green colour was observed.	
Substance E is a white crystalline solid. It was dissolved in water.	The solution was blue and was warm to the touch.	
A small amount of E was dissolved in distilled water. The solution was acidified with hydrochloric acid and dilute barium chloride added.	A whitish-blue precipitate was formed which was insoluble in excess acid.	
	Name of E:	
	Formula of E:	

2.

TEST	OBSERVATION	INFERENCE
Pure substance F was heated strongly, on a platinum wire, in a bunsen flame.	A whitish flame colour was observed.	
Substance F was a pale green crystalline solid soluble in water. Dilute sodium hydroxide was added to the solution.	A green precipitate was formed. This turned brownish on standing.	
A small amount of F was heated strongly in a borosilicate glass tube.	A colourless liquid formed on the cool part of the tube, a reddish brown solid was formed and acidic, brown, fumes were evolved.	
A small amount of F was heated with Devarda's alloy and sodium hydroxide solution.	A pungent smelling gas was evolved which turned damp red litmus blue. The gas gave white fumes with the vapour from concentrated hydrochloric acid.	
	Name of F:	
	Formula of F:	

Test 46: Theme: Practical work - Organic Analysis - 1

Complete the inference columns and write a conclusion, name &/or formula as required.

1.

TEST	OBSERVATION	INFERENCE
Pure organic substance A was combusted on a crucible lid.	It burnt with a smokey flame.	
Substance A was shaken with some bromine water.	The bromine water was decolourised.	
Substance A was shaken with dilute potassium permanganate solution.	The permanganate solution was decolourised.	
Quantitative analysis showed that A reacted with hydrogen gas.	1 mole of A reacted with 1 mole of hydrogen gas	
	Conclusion:	

2.

TEST	OBSERVATION	INFERENCE
Substance B is a colourless liquid with a pleasant odour. A small amount was burnt on a crucible lid.	It burnt with a clean flame. The odour is alcoholic.	
Substance B was added to water.	B mixed with water in all proportions. The solution became warm.	
Substance B was added to Brady's reagent.	There was no change (no apparent reaction).	
Substance B was mixed with potassium dichromate in dilute sulphuric acid. The mixture was warmed in hot water.	Over a period of a few minutes the mixture changed colour from yellow-orange to green. A faint smell of over ripe apples was detected.	
The above test (test 4) was repeated and the vapours drawn into a teat pipette and expelled into Tollens reagent.	A black precipitate and silver mirror was formed.	
	Name of B:	
	Formula of B:	

Test 47: Theme: Practical work - Organic Analysis - 2

Complete the inference columns and write a conclusion, name &/or formula as required.

1.

TEST	OBSERVATION	INFERENCE
Substance C is a colourless liquid. It was added to water.	Substance C mixes with water in all proportions. It has a vinegary odour. The solution turns blue litmus red.	
A few drops of C was added to 10% sodium carbonate solution.	Effervescence occurred. The gas evolved turned lime water milky.	
Six drops of C were added to 1 cm ³ of ethanol in a test tube. A drop of concentrated sulphuric acid was added and the mixture warmed in warm water.	When the contents of the tube were poured into water in a beaker a strong fruity odour was apparent.	
	Write a name for C:	
	Write a formula for C:	

2.

TEST	OBSERVATION	INFERENCE
Substance D is yellowish, oily, liquid. It was added to water in a test tube.	Substance did not appear to dissolve in the water however, pH paper indicated that the aqueous solution was slightly alkaline.	
A drop of D was burned on a crucible lid.	A luminous sooty flame was produced.	
A few drops of D was added to 1 cm ³ of ethanoic anhydride (acetic anhydride). The mixture was shaken in a test tube.	A white precipitate was formed. This was filtered off, purified, and its melting point measured (mp 114°C).	
A few drops of D was shaken with bromine water.	A yellowish precipitate was formed.	
	Write a name for D:	
	Write a formula for D:	

Test 48: Theme: Practical work - Organic analysis - 3

Complete the inference column and the conclusions.

1.

TEST	OBSERVATION	INFERENCE
Substance E is a white crystalline solid (mp 137°C). Test solubility in water.	It is soluble in <u>hot</u> water. Giving an acidic solution.	
Burn E on a crucible lid.	Very smokey, luminous flame.	
Following the Lassaigne procedure E was fused with sodium and then digested in distilled water. The filtrate was tested for nitrogen, halogens, & sulphur.	The tests indicated that, apart from carbon, hydrogen and oxygen the only other element present was chlorine.	
Infra-red spectroscopy	Sharp, medium sized peaks just above 3000 cm ⁻¹ . Very broad absorption 3500-2700 cm ⁻¹ .	
Substance E was refluxed with ethanol and a few drops of concentrated sulphuric acid. The reaction mixture was distilled and a pleasant smelling liquid was isolated.	The distillate boiled at 243°C and gave a positive ester test (alcoholic hydroxylamine hydrochloride, methanolic KOH, boil, acidify with HCl and add a drop or two of iron(III) chloride - deep red or purple indicates ester).	
	Name of E:	
	Formula of E:	

2.

TEST	OBSERVATION	INFERENCE
Substance F: Very cautiously smell. Test solubility in water.	It is highly volatile. It has a very pungent odour. Soluble in water.	
Infra-red spectroscopy.	Intense absorption at 1727 cm ⁻¹ . Medium absorption at 2733 cm ⁻¹ . Medium to strong absorption at 2846 cm ⁻¹ .	
Nuclear magnetic resonance spectroscopy.	Singlet peak at 2.20 ppm (integration = 3 units) Singlet peak at 9.79 ppm (integration = 1 unit)	
	Name of F:	
	Formula of F:	

Test 49: Theme: Practical work - Organic analysis - 4

Complete the inference column and identify G.

1.

TEST	OBSERVATION	INFERENCE
Substance G is a colourless liquid. Test its solubility in water.	Slightly soluble in water giving a neutral solution. G has an odour of roses.	
Burn a few drops of G on a crucible lid.	It burns with a very smokey flame.	
Apply Lassaigne test.	G does not contain halogens, nitrogen or sulphur.	
Apply quantitative C, H & O analysis. Also mass spectrometry analysis.	C = 78.69%, H = 8.19%, O = 13.11%. Relative molecular mass = 122	
Esterify with ethanoic acid and a few drops of concentrated sulphuric acid.	Following work up and distillation a pleasant smelling, colourless, liquid was obtained boiling at 232°C.	
Infra-red spectroscopy.	Broad intense peak centred at 3500 cm ⁻¹ . Strong peaks just below 3000 cm ⁻¹ . Sharp peaks just above 3000 cm ⁻¹ .	
Pmr spectroscopy.	Singlet at 7.2 ppm (integration = 5 units) Triplets at 2.8 and 3.8 ppm (integration = both sets 2 units each). Singlet at 2.1 ppm (integration = 1 unit).	
	Name of G:	
	Formula of G:	

Test 50: Theme: Practical work - Organic analysis - 5

Complete the inference column and identify H.

1.

TEST	OBSERVATION	INFERENCE
Substance H is a colourless liquid (bp 185°C). Solubility in water.	Partially soluble in water. The solution was alkaline to pH paper.	
Burn on a crucible lid.	Very smokey, luminous, flame.	
Lassaigne analysis.	Positive result for nitrogen. Negative results for halogen and sulphur.	
Shake a few drops of H with 1 cm ³ of benzoyl chloride and 2 cm ³ of cold, dilute, sodium hydroxide.	Grey-white precipitate.	
Add a few drops of H to an ice cold solution of nitrous acid.	Efforescence occurred. The gas was colourless & neutral.	
Quantitative analysis.	Carbon = 78.50%, H = 8.41%, N = 13.08%.	
Mass spectrometry:	Mol. Ion peak = 107 Base peak = 106	
Infra-red spectroscopy:	3400 cm ⁻¹ 3030 - 3090 cm ⁻¹ 2860 - 2930 cm ⁻¹	
Proton magnetic resonance spectroscopy:	Singlet at 1.50 ppm (integration = 2 units) Singlet at 3.80 ppm (integration = 2 units) Peaks centred at 7.35 ppm (Integration = 5 units)	
	Name of H:	
	Formula of H:	

Test 51: Theme: Practical work - Organic analysis - 6

Complete the inference column and identify I.

1.

TEST	OBSERVATION	INFERENCE
Substance I is a colourless liquid. Bp 173°C. Test solubility in water.	I has an ester-like odour. It is insoluble in water.	
Burn I on a crucible lid.	I burns readily with a relatively clean, luminous, flame.	
Test I with Brady's reagent.	Yellow-orange precipitate. Following recrystallisation and drying it melted at 84°C.	
Add a few drops of I to bromine solution and shake.	Bromine solution was decolourised (red-brown to colourless).	
Quantitative analysis. Mass spectrometry.	C = 76.19%, H = 11.11%, O = 12.70%. Molecular ion peak = 126.	
Infra red spectroscopy	Intense absorption at 1720 cm ⁻¹ . Intense absorption just below 3000 cm ⁻¹ . No absorption immediately above 3000 cm ⁻¹ .	
Proton magnetic resonance	Peak at 1.65 ppm (integration = 6 units). Singlet at 2.13 ppm (integration = 3 units). Peak at 2.25 ppm (integration = 2 units). Peak at 2.46 ppm (integration = 2 units). Peak at 5.07 ppm (integration = 1 unit).	
	Name of compound I: Formula of compound I:	